Solutions to:
Gas Law Homework Problem Set
Chemistry 145, Chapter 12

1. The volume of a bicycle tire is 1.35 liters and the manufacturer recommends a tire pressure of 125 PSI.

a. If you want the bicycle tire to have the correct pressure at 20.0 °C, what volume of air is required at STP?

Information given in question:

$V := 1.35 \cdot liter$

$P := 125 \cdot psi$ $\qquad P = 8.618 \times 10^5\ Pa$

$T := (273.15 + 20) \cdot K$ $\qquad T = 293.15\ K$

Note: you may work the problem using any pressure units, BUT you must use the same units for standard pressure and for the tire pressure.

Conditions at STP (Standard Temperature and Pressure):

$P_{STP} := 1 \cdot atm$ $\qquad P_{STP} = 1.013 \times 10^5\ Pa$

$T_{STP} := 273.15 \cdot K$ $\qquad T_{STP} = 273.15\ K$

Mathematical relationship (the combined gas law):

$$\frac{P_1 \cdot V_1}{T_1} = \frac{P_2 \cdot V_2}{T_2}$$

Rearranges to

$$V_1 = P_2 \cdot \frac{V_2}{(T_2 \cdot P_1)} \cdot T_1$$

Substitute in variables for this problem

$$V_{STP} := P \cdot \frac{V}{(T \cdot P_{STP})} \cdot T_{STP}$$

$$V_{STP} = 10.699\ liter$$

b. If you fill the tire with nitrogen, what is the mass of the gas?

Calculate the moles of gas using the ideal gas law:

$$P \cdot V = n \cdot R \cdot T \qquad R := 8.314510 \cdot joule \cdot K^{-1} \cdot mole^{-1}$$

$$n := \frac{P \cdot V}{R \cdot T}$$

$$n = 0.477\ mole$$

Calculate the mass of the nitrogen:

$$MW := (2 \cdot 14.00674) \cdot gm \cdot mole^{-1}$$

$$mass := n \cdot MW$$

$$mass = 13.372\ gm$$

c. If you fill the tire with compressed gas from a 100.0 mL cylinder at 1.50×10^7 Pa, what is the final pressure in the tire? Will it explode?

Information given in question:

$$V_{cyl} := 100.0 \cdot mL$$

$$P_{cyl} := 1.50 \cdot 10^7 \cdot Pa \qquad P_{cyl} = 2.176 \times 10^3\ psi$$ (*VERY high pressure*)

Mathematical relationship (the combined gas law):

$$\frac{P_1 \cdot V_1}{T_1} = \frac{P_2 \cdot V_2}{T_2}$$

Since the temperature is constant this simplifies to Boyle's Law

$$P_1 \cdot V_1 = P_2 \cdot V_2$$

Which rearranges to:

$$P_1 = \frac{P_2 \cdot V_2}{V_1}$$

Substituting in variables for this problem gives:

$$P := \frac{P_{cyl} \cdot V_{cyl}}{V}$$

$$P = 1.111 \times 10^6\ Pa$$

$$P = 161.153\ psi$$

(*Looks like you are pushing your luck and this tire just might pop. Better be careful*)

d. You fill the bicycle tire to 125 PSI on a cold December day (-22 °F), and leave it until a hot day in July (101 °F). What is the pressure of the tire (assuming that it does not leak, does not change volume, and does not burst).

Information given in question:

$$P_{dec} := 125 \cdot psi$$

$$T_{dec} := \left[273.15 + \left[\frac{5}{9} \cdot (-22 - 32)\right]\right] \cdot K \qquad T_{dec} = 243.15\,K$$

$$T_{jul} := \left[273.15 + \left[\frac{5}{9} \cdot (101 - 32)\right]\right] \cdot K \qquad T_{jul} = 311.483\,K$$

Mathematical relationship (the combined gas law):

$$\frac{P_1 \cdot V_1}{T_1} = \frac{P_2 \cdot V_2}{T_2}$$

Since the volume is constant this simplifies to Charles's Law:

$$\frac{P_1}{T_1} = \frac{P_2}{T_2}$$

Rearrange and substitute in variables from above:

$$P_{jul} := \frac{P_{dec}}{T_{dec}} \cdot T_{jul}$$

$$P_{jul} = 160.129\,psi$$ (*Looks like you may have a problem here)*

2. A fire extinguisher with a volume of 3.0 liters is filled with 150.0 grams of CO_2. Assuming that it is an ideal gas, what is the pressure at 20.0 ^oC?

Information given in problem:

$V_{extinguisher} := 3.0 \cdot liter$

$T := (273.15 + 20.0) \cdot K$

$CO2_{mass} := 150.0 \cdot gm$

To solve this using the ideal gas law, we need to know how many moles of CO_2.

$CO2_{MW} := (12.0107 + 2 \cdot 15.9994) \cdot gm \cdot mole^{-1}$ $\quad CO2_{MW} = 44.01\ gm \cdot mole^{-1}$

$CO2 := \frac{CO2_{mass}}{CO2_{MW}}$

$CO2 = 3.408\ mole$

The ideal gas law:

$P \cdot V = n \cdot R \cdot T$

$n := CO2$

$R = 8.315\ joule \cdot K^{-1} \cdot mole^{-1}$

$T = 293.15\ K$

$V := V_{extinguisher}$

$P := \frac{n \cdot R \cdot T}{V}$

$P = 2.769 \times 10^6\ Pa$

$P = 27.33\ atm$

$P = 2.077 \times 10^4\ torr$

(*Note: Be certain to use units for all values so that they will cancel. Using this value of R (with SI units), pressure should be in Pa. You must always use absolute temperature units for gas law calculations. This gives volume in m^3 (SI units).)*

The volume at 1.0 atm and 25 ^oC after the CO_2 is released from the fire extinguisher?

$T := (273.15 + 25) \cdot K$ $\quad P := 1 \cdot atm$ $\quad P = 1.013 \times 10^5\ Pa$

$V := \frac{n \cdot R \cdot T}{P}$

$V = 0.083\ m^3$

$V = 83.387\ liter$

(*Note: Be certain to use units for all values so that they will cancel. Using this value of R (with SI units), pressure should be in Pa. You must always use absolute temperature units for gas law calculations. This gives volume in m^3 (SI units).)*

3. A diver is using a NITROX mixture (oxygen enriched air to reduce risk of decompression illness and increase bottom time) to breath underwater. If the mixture is 36% oxygen and 64% nitrogen, what is the partial pressure of oxygen (Pa) when the diver is at

a. sea level (1.0 atm)

$P := 1.0 \cdot atm$ $\quad P = 1.013 \times 10^5\ Pa$

$P_{O2} := P \cdot 36 \cdot \%$ $\quad P_{O2} = 0.36\,P$ $\quad P_{O2} = 3.648 \times 10^4\ Pa$ $\quad P_{O2} = 0.36\ atm$

b. 33.9 feet (2.0 atm)

$P := 2.0 \cdot atm$ $\quad P = 2.026 \times 10^5\ Pa$

$P_{O2} := P \cdot 36 \cdot \%$ $\quad P_{O2} = 0.36\,P$ $\quad P_{O2} = 7.295 \times 10^4\ Pa$ $\quad P_{O2} = 0.72\ atm$

c. 100 ft (3.9 atm)

$P := 3.9 \cdot atm$ $\quad P = 3.952 \times 10^5\ Pa$

$P_{O2} := P \cdot 36 \cdot \%$ $\quad P_{O2} = 0.36\,P$ $\quad P_{O2} = 1.423 \times 10^5\ Pa$ $\quad P_{O2} = 1.404\ atm$

a. 500 ft (15.8 atm)

$P := 15.8 \cdot atm$ $\quad P = 1.601 \times 10^6\ Pa$

$P_{O2} := P \cdot 36 \cdot \%$ $\quad P_{O2} = 0.36\,P$ $\quad P_{O2} = 5.763 \times 10^5\ Pa$ $\quad P_{O2} = 5.688\ atm$

4. A balloon used for sampling stratospheric ozone is filled with 150.0 kg of He. What is the volume of the balloon when,

First we need to determine the number of moles of He:

$He_{mass} := 150.0 \cdot kg$ $\quad He_{mass} = 1.5 \times 10^5\ gm$

$He_{MW} := 4.002602 \cdot gm \cdot mole^{-1}$

$He := \dfrac{He_{mass}}{He_{MW}}$ $\quad He = 3.748 \times 10^4\ mole$

$n := He$

a. The balloon starts at sea level at a research station in Antarctica where the barometric pressure is 755 mmHg and the temperature is -25 °C.

$mmHg := 1 \cdot torr$

$T := (273.15 - 25) \cdot K$

$P := 755 \cdot mmHg$ $\qquad P = 1.007 \times 10^5\ Pa$

(*Note: 1 millimeter of mercury (mmHg) is equal to 1 torr. This pressure unit was widely used in the US, but it is not recommended for SI units. The preferred unit is the pascal (Pa).*)

The ideal gas law:

$$P \cdot V = n \cdot R \cdot T$$

rearranges to

$$V := \frac{n \cdot R \cdot T}{P} \qquad V = 768.159\ m^3 \qquad V = 7.682 \times 10^5\ liter$$

b. The balloon rises to 10,000 ft (3048 m, the height of a medium size mountain) where the instruments report that the temperature is -50 °C and the pressure is 6.368×10^4 Pa.

Information given in the problem:

$T := (273.15 - 50) \cdot K \qquad P := 6.368 \cdot 10^4 \cdot Pa$

The ideal gas law:

$$V := \frac{n \cdot R \cdot T}{P}$$

$$V = 1.092 \times 10^3\ m^3$$

c. The balloon continues to rise, at 29,028 ft (8,848 m, the height of Mt. Everest and about typical cruising altitude for a jet aircraft) the temperature is -70 °C and the pressure is 2.30×10^4 Pa.

Information given in the problem:

$T := (273.15 - 70) \cdot K \qquad P := 2.30 \cdot 10^4 \cdot Pa$

The ideal gas law:

$$V := \frac{n \cdot R \cdot T}{P}$$

$$V = 2.752 \times 10^3\ m^3$$

d. The balloon enters the stratosphere, at 65,000 ft (20,000 m, cruising altitude for a U2 spy plane) the temperature is -50 ^{o}C and the pressure is 3.68x10^{3} Pa.

Information given in the problem:

$$T := (273.15 - 70)\cdot K \qquad P := 3.68\cdot 10^{3}\cdot Pa$$

The ideal gas law:

$$V := \frac{n\cdot R\cdot T}{P}$$

$$V = 1.72 \times 10^{4}\, m^{3}$$

e. The balloon reaches its maximum altitude of 100,000 ft (20500 m) the temperature is -3 ^{o}C and the pressure is 616 Pa.

Information given in the problem:

$$T := (273.15 - 3)\cdot K \qquad P := 616\cdot Pa$$

The ideal gas law:

$$V := \frac{n\cdot R\cdot T}{P}$$

$$V = 1.367 \times 10^{5}\, m^{3}$$

5. I recieved the following question from Larry Stratton (larstrat@centurytel.net), a retired civil engineer. Assume a blocked length of pipe with a contained volume of 10 ft^3. Assume air is injected until a pressure gauge reading of 3.5 psi is stabilized. Then assume an air leak at the rate of 0.003 ft^3/min. How long will it take for the air pressure to drop to 2.5 psi gauge reading? In the real world of low pressure air testing for sewer pipes, there are factors not considered here, i.e., type of pipe etc., but this will do as an exercise. larstrat@centurytel.net

$P_{initial} := 3.5 \cdot psi$

$P_{final} := 2.5 \cdot psi$

$P_{drop} := P_{initial} - P_{final}$

$P_{drop} = 1\ psi$ $\qquad$ $P_{drop} = 6.895 \times 10^3\ Pa$

$V := 10 \cdot ft^3$ $\qquad$ $T := 273.15 \cdot K$ $\qquad$ Assume standard temp

$n := \frac{P_{drop} \cdot V}{R \cdot T}$ $\qquad$ $n = 0.86\ mole$ $\qquad$ moles of gas that escapes

Next find volume of air that escapes at STP

$P := 10^5 \cdot Pa$

$V := \frac{n \cdot R \cdot T}{P}$

$V = 0.02\ m^3$ $\qquad$ $V = 0.689\ ft^3$

Time for the air to leak

$Leak := 0.003 \cdot ft^3 \cdot min^{-1}$

$time := \frac{V}{Leak}$ $\qquad$ $time = 229.825\ min$

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